PLI 36

Thursday July 21, 2016

$\frac{0.010 \text{ mol NaOH}}{1 \text{ L solution}} =$ [0.010 NaOH]

Chapter 17

$pOH = -\log(0.010) = 2$

$pH = 14 - 2 = \boxed{12}$

1. A 1.0-L buffer solution contains 0.100 mol acetic acid and 0.100 mol sodium acetate. The value of *K*a for acetic acid is 1.8 x 10^{-5}. Calculate the new pH after adding 0.010 mol of solid NaOH to the buffer. For comparison, calculate the pH after adding 0.010 mol of solid NaOH to 1.0 L of pure water.

1)

	$HC_2H_3O_2$ +	NaOH $\longrightarrow$	$NaC_2H_3O_2$ +	H_2O
initial :	0.100 mol	0	0.100 mol	
added :	− 0.010 mol	0.010 mol	+ 0.010 mol	
post rxn :	0.090 mol	0 mol	0.110 mol	

2) plug into H-H:

$pH = pka + \log\frac{[C_2H_3O_2]}{[HC_2H_3O_2]} = 4.74 + \log\frac{(0.110)}{(0.090)} = \boxed{4.83}$

2. A 500.0-mL buffer solution is 0.100 M in HNO_2 and 0.150 M in KNO_2. Determine if each addition would exceed the capacity of the buffer to neutralize it.

→ exceed buffer capacity if add acid/base beyond the amount the buffer can neutralize. In this buffer, there are

$0.500 L \times \frac{0.100 \text{ mol } HNO_2}{L} = 0.050$ mol acid

$0.500 L \times \frac{0.150 \text{ mol } KNO_2}{L} = 0.075$ mol base

a. 2.50 g NaOH

$2.50g \left(\frac{1 \text{mol NaOH}}{40g}\right)\left(\frac{1 \text{mol } OH}{1 \text{mol NaOH}}\right) = 0.0625$ mol OH

mol OH^- > mol HNO_2 so buffer won't work.

b. 1.25 g HBr

$1.25g \text{ HBr}\left(\frac{1 \text{mol HBr}}{80.9g}\right)\left(\frac{1 \text{mol } H^+}{1 \text{mol Br}}\right) = 0.015$ mol H^+

mol H^+ < mol NO_2^-, so buffer works

c. 3.60 g KOH

$3.60 g \text{ KOH}\left(\frac{1 \text{mol KOH}}{56.11g}\right)\left(\frac{1 \text{mol } OH}{1 \text{mol KOH}}\right) = 0.064$ mol OH

mol OH^- > mol HNO_2, buffer won't work

d. 1.35 g HI

$1.35 g\text{ HI}\left(\frac{1 \text{mol HI}}{127.91}\right)\left(\frac{1 \text{mol } H^+}{1 \text{mol HI}}\right) = 0.011$ mol H^+ → mol H^+ < mol NO_2^-, buffer works

3. A chemist has synthesized a monoprotic weak acid and wants to determine its *K*a value. To do so, the chemist dissolves 2.00 millimoles of the solid acid in 100.0 mL of water and titrates the resulting solution with 0.0500 M NaOH. After 20.0 mL NaOH has been added, the pH is 6.00. What is the *K*a value for the acid?

1)

	HA +	NaOH $\longrightarrow$	NaA +	H_2O
initial :	0.002 mol	0	0	
added :	−0.001 mol	0.001 mol	+0.001 mol	
post rxn :	0.001 mol	0	0.001 mol	

2) HH:

$pH = pka + \log\left(\frac{A^-}{HA}\right)$

$6.00 = pka + \log\left(\frac{0.001}{0.001}\right)$

$pka = 6.00$ $\boxed{Ka = 10^{-6}}$

4. Find the pH at each of the following points in the titration of 25 mL of 0.30 M HF with 0.30 M NaOH. The Ka HF=7.2 x 10^{-4}

   a. Before adding NaOH

	HF +	H_2O ⇌	F^- +	H_3O^+
I	0.30M		0	0
C	−x		+x	+x
E	0.30 − x		x	x

$Ka = \frac{x^2}{(0.30 - \cancel{x})}$ ← can ignore

$(7.2 \times 10^{-4})(0.30) = x^2$

$x = 0.0147$

$pH = -\log(0.0147) = \boxed{1.83}$

   b. After adding 10.00 mL of NaOH

	HF +	NaOH →	NaF +	H_2O
initial:	0.0075 mol	0	0	
added:	−0.0030mol	0.0030mol	0.0030mol	
end:	0.0045 mol	0	0.0030 mol	

in a buffer, use HH:

$pH = pka + \log\left(\frac{B}{A}\right)$

$pH = 3.14 + \log\left(\frac{0.0030}{0.0045}\right) = \boxed{2.97}$

   c. At ½ equivalence point

$pH = pka$

$\boxed{pH = 3.14}$

   d. At the equivalence point

1) to titrate 0.0075 mol HF, need 0.0075 mol NaOH; if [NaOH] was 0.300M, then added 25 mL of NaOH.

2) so created 0.0075 mol F^- in 0.050L sln → $[F^-] = \frac{0.0075}{0.050} = 0.15M$

3) set up ICE table

F^- +	H_2O ⇌	HF +	OH^-
0.15		0	0
−x		+x	+x
0.15 − x		x	x

4) solve using kb

$1.4 \times 10^{-11} = \frac{x^2}{0.15 - \cancel{x}}$

$x = 1.4 \times 10^{-6}$

$pOH = -\log(1.4 \times 10^{-6}) = 5.84$ $\boxed{pH = 8.16}$

   e. After adding 28.00 mL of NaOH

only care about [OH⁻] from NaOH, because it is much more basic than F⁻.

0.003L NaOH $\left(\frac{0.3 \text{ mol}}{L}\right)$ = **0.0009 mol** NaOH

$\frac{\textbf{0.0009 mol} \text{ NaOH}}{0.053L \text{ solution}}$

[OH⁻] = **0.017 M**

pOH = **1.77**

pH = **12.23**

5. A solution is prepared by adding 750.0 mL of 4.00 x 10^{-3} M Ce(NO_3)$_3$ to 300.0 mL of 2.00 x 10^{-2} M KIO_3. Will Ce(IO_3)$_3$ (K_{sp} = 1.9 x 10^{-10}) precipitate from this solution?

1) find moles Ce^{3+} and IO_3^- you had in original solutions and calculate molarities in new solution

$0.750L\ Ce^{3+}\text{ solution} \left(\frac{4 \times 10^{-3} \text{ mol}}{1L}\right) = 0.00300 \text{ mol } Ce^{3+}$

$0.300L\ IO_3^-\text{ solution} \left(\frac{2 \times 10^{-2} \text{ mol}}{1L}\right) = 0.00600 \text{ mol } IO_3^-$

→

2) new concentrations

$$\frac{0.00300 \text{ mol } Ce^{3+}}{(0.750L + 0.300L)} = 0.00286 \text{ M } Ce^{3+}$$

$$\frac{0.00600 \text{ mol } IO_3^-}{(0.750L + 0.300L)} = 0.00571 \text{ M } IO_3^-$$

3) now its a Q_{sp} vs. K_{sp} problem

$$Ce(IO_3)_3\ (s) \rightleftharpoons Ce^{3+}(aq) + 3IO_3^-(aq)$$

$$Q = [Ce^{3+}] \times [IO_3^-]^3$$

$$Q = (0.00286)(0.00571^3) = 5.33 \times 10^{-10}$$

$Q_{sp} > K_{sp}$

$5.33 \times 10^{-10} > 1.9 \times 10^{-10}$, means we have too many products, so rxn shifts to left and $Ce(IO_3)_3$ precipitates out.